SOLUTIONS

RATE LAWS WORKSHEET I

1. Using the experimental data provided, determine the order of reaction with respect to each reactant, write the rate law, determine the overall order of the reaction, and calculate the rate law constant, k.

$S_2O_8^{2-}$ (aq) + 3 I^- (aq) → 2 SO_4^{2-} (aq) + I_3^- (aq)

Experiment	Initial Concentration (mol/L)		Initial Rate (mol/L•s)
	$S_2O_8^{2-}$	I^-	
1	0.15	0.21	1.14
2	0.22	0.21	1.70
3	0.22	0.12	0.98

$S_2O_8^{2-}$ 1st order

I^- 1st order

Rate = $k[S_2O_8^{2-}][I^-]$ ← Rate Law

$1.14 = k(0.15)(0.21)$

$k = 36.2\ L/mol \cdot s$ ← Rate Law Constant.

2. Using the experimental data provided, determine the order of reaction with respect to each reactant, the rate law equation, the overall order of reaction, and calculate the rate law constant, k.

CO (g) + Cl_2 (g) → $COCl_2$ (g)

Experiment	Initial Concentration (mol/L)		Initial Rate (mol/L•s)
	CO	Cl	
1	0.12	0.20	0.121
2	0.24	0.20	0.241
3	0.12	0.40	0.483

CO - 1st order

Cl - 2nd order

Rate = $k[CO][Cl^-]^2$

$0.121 = k(0.12)(0.20)^2$

$k = 25.2\ L^2/mol^2 \cdot s$

3. Using the experimental data provided, determine the order of reaction with respect to each reactant, the rate law equation, the overall order of reaction, and calculate the rate law constant, k. Use the data to predict the reaction rate for Experiment 4.

2 ICl (g) + H_2 (g) → I_2 (aq) + 2 HCl (g)

Experiment	Initial Concentration (mol/L)		Initial Rate (mol/L•s)
	ICl	H_2	
1	1.5	1.5	3.7×10^{-7}
2	3.0	1.5	7.4×10^{-7}
3	3.0	4.5	2.2×10^{-6}
4	4.7	2.7	?

ICl - 1st order

H_2 - 1st order

Rate = $k[ICl][H_2]$

$3.7 \times 10^{-7} = k(1.5)(1.5)$

$k = 1.64 \times 10^{-4}\ L/mol \cdot s$

Rate = $(1.64 \times 10^{-7})(4.7)(2.7)$

Rate = $2.1 \times 10^{-6}\ mol/L \cdot s$

4. Using the experimental data provided, determine the order of reaction with respect to each reactant, the rate law equation, the overall order of reaction, and calculate the rate law constant, k. Use the data to predict the reaction rate for Experiment 4.

$NO_2\ (g) + O_3\ (g) \rightarrow NO_3\ (g) + O_2\ (g)$

Experiment	Initial Concentration (mol/L)		Initial Rate (mol/L•s)
	NO_2	O_3	
1	0.21	0.70	6.3
2	x1.8 0.21	1.39 x2	x1.8 12.5 x2
3	0.38	0.70	11.4
4	0.66	0.18	?

NO_2 - 1st order
O_3 - 1st order

Rate = $k[NO_2][O_3]$

$6.3 = k(0.21)(0.70)$

$k = 42.9$ L/mol·s

Rate = $42.9(0.66)(0.18)$

Rate = 5.09 mol/L·s

5. The reduction of bromate ions, BrO_3^-, by bromide ions in acidic solution has a rate law $R = k[BrO_3^-][Br^-][H^+]^2$.
   a. What are the orders with respect to the reactants?
   b. What is the overall order?

BrO_3^- - 1st order
Br^- - 1st order
H^+ - 2nd order

Overall 4th order.

6. The reaction between bromomethane and hydroxide ion in aqueous solution is first order with respect to bromomethane, and second order overall. Write the rate law.

Rate = $k[CH_3Br][OH^-]^2$

↑ formula for bromomethane (you will learn about this later)